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BALLOONS

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Here, a balloon serves as an analogy for a bubble.

Skin Tension;

When you begin to blow up a balloon, it's difficult to inflate. As it grows, it gets easier. There are a number of reasons why this is so, but for these purposes, say that it is due to skin tension. The rubber skin exerts an inward force over the surface, causing the pressure inside the balloon to be higher than external pressure (resulting in a "pop" when pierced). Skin tension varies inversely with the radius of a balloon. This causes the pressure inside a smaller balloon to be greater than a larger--resulting in the decrease in inflation resistance as the balloon grows (Fig B2).

Of course, it never becomes effortless to blow up a balloon (unless it pops). So, we see that there is an important effect for large bubbles also. This is the elasticity, or compliance (springy-ness) of displaced tissue surrounding the bubble. This volume-dependent effect serves to aid shrinking large bubbles just as skin tension helps collapse smaller bubbles.

Diffusion

If a balloon inflated with helium is left for a few days, it will deflate. This is obviously due to the helium "leaking," or diffusing, through the skin. The balloon's skin is permeable to He, allowing outward flow to occur. Furthermore, it is the pressure gradient G from inside the balloon to outside the balloon that drives the gas outward. Even if the balloon were immersed in pure helium, the balloon would still collapse because its internal pressure would be higher than external due to the skin tension (this is similar to the concept of the oxygen window in living creatures).

Internal Pressure

Consider the following: Two identical balloons are inflated, one more than the other. A closed valve connects the large and the small balloons.

Question: When the valve is opened, what happens?

Answer: A reasonable guess is that the pressures would equalize on either side of the valve, with gas flowing from high pressure to low. Recalling that a smaller balloon has higher internal pressure than a larger, we see the gas must flow from the small to the large balloon. Counter to the naive idea that the two balloons would end up with identical sizes, the smaller balloon deflates as the larger balloon inflates.

The important points are that a bubble's skin causes its internal pressure to be greater than

ambient, and the pressure inside small bubbles is greater than larger bubbles. Other physical effects such as Boyle's law and tissue compression also affect the internal pressure of bubbles. All of this is important for developing bubble decompression tables because ascents are limited by the pressure difference between bubbles and surrounding tissue tension. We need to account for the physical effects that influence bubble pressure in order to quantify ascent criteria. This awareness also allows us to use sound physical principles to control the inflation of bubbles. These guidelines hold generally, regardless of whether we are talking about gelatin, balloons, beer or humans.
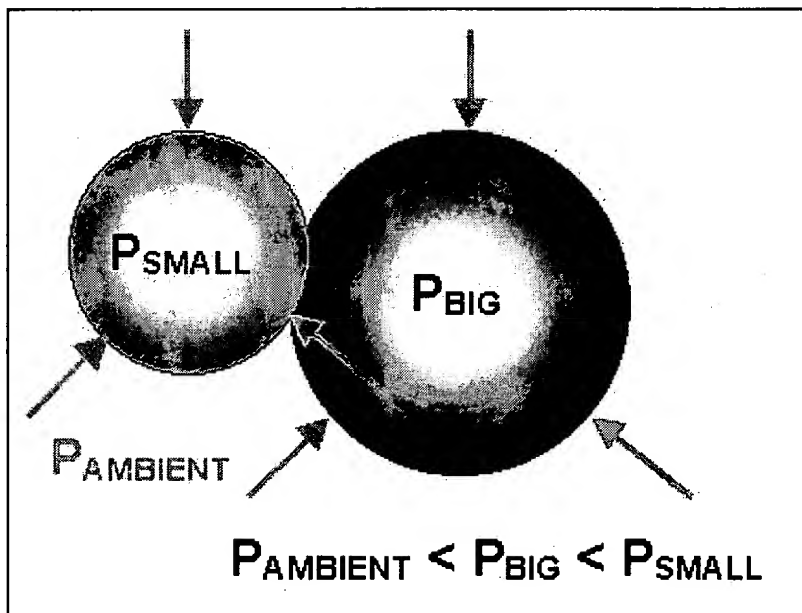

Fig. B2. The pressure inside balloons, as with bubbles, is greater than ambient. Because the Laplacian skin tension varies as 1/radius, the pressure inside the smaller balloon is larger than the pressure inside the big balloon.

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What is the Pressure inside the balloon?

Questions Questions!

The more you work on a project the more questions you have. In trying to better anticipate (in other words: predict) possible problems and successes we have had to address questions like: How much helium will we need? What size balloon do we need? How high can it go? What will happen to it, as it gets high up? What can we use for a fill tube? How can we hold on to the balloon while it is filling?

All questions lead to other questions and most of these brought up the question of: What is the pressure inside the balloon and how does it change with the size?

As you know from blowing up party balloons, you have to blow really hard at first and then it gets easy until just before it bursts. This would suggest that the pressure inside changes dramatically depending on the size of the balloon. Is this the same as for our huge meteorological balloons? How much of a pressure difference is this anyway?

This is the kind of fundamental question that we can directly measure.

Planning the Setup

We try to keep this simple. We have instruments that can measure pressure and after trying it out we can stretch the neck of a balloon enough to get a barometric sensor inside.

We need a source of gas. No it doesn't have to be helium. What about air? OK but it has to be under some pressure and capable of very large volume. My lung-power is out. The lab has a spigot marked "air".

The barometer will record the pressure inside the balloon twice a second and we can fill at a constant flow rate. However, we want to relate this to the size of the balloon.

It is very hard to measure diameter. However, we know that both diameter and volume are easily calculated from circumference. We try a tape measure and it is too stiff to go around a balloon. We take a string and mark it off in feet instead.

The Experiment

We recruit extra people and clear an area for space to work and to remove any sharp edges. One person holds the balloon, one holds the circumference string, & one has a stopwatch and a data table. The stopwatch and barometer are started and the room air pressure data is collected for a few seconds before we start the air flowing into the balloon. When the balloon first gets rounded out we record the time and circumference. As the balloon's circumference increases to each additional footmark, we record the time.

It is getting longer and longer between each new foot of circumference. This makes sense because the volume is going up as the cube of the diameter. How long do we continue this? Not much appears to be happening but we are nowhere near the bursting point. We decide to inflate well past the largest size we expect to need on the ground and then extrapolate for the expected size gain as the balloon rises to our 1000 ft standard height. The last foot seems to

take forever. We mark the time, shutoff the air, open the balloon neck, and it seems to take forever to deflate also.

However, when we examine the data it looks good. There is definitely a relationship between the pressure and time (inflation). But what does it mean?
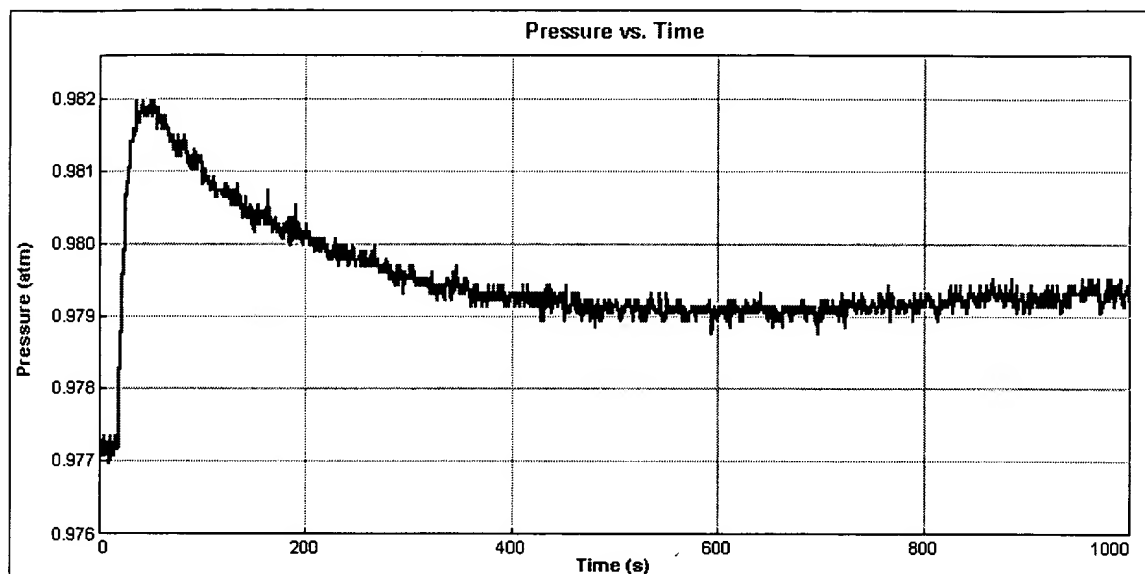
Time	circumference in ft.	radius in ft.	diameter in ft.	volume in ft^3
0	-	-	-	-
37	5.25	0.84	1.67	2.44
59	6	0.95	1.91	3.65
90	7	1.11	2.23	5.79
130	8	1.27	2.55	8.65
186	9	1.43	2.86	12.31
240	10	1.59	3.18	16.89
307	11	1.75	3.50	22.48
405	12	1.91	3.82	29.18
519	13	2.07	4.14	37.10
644	14	2.23	4.46	46.34
839	15	2.39	4.77	56.99
921	15.5	2.47	4.93	62.88
990	16	2.55	5.09	69.17

Analysis

Looking at the graph, note that the horizontal scale is in seconds. 1000 seconds is about 17 minutes. Note the vertical scale. It starts at 0.976 atm instead of zero. This means that the vertical changes are exaggerated. The graph starts off fairly constant at just above 0.977 atm. This must be the room pressure in the lab about 98% of average standard air pressure.

The pressure rapidly goes up to a high of 0.982 before it drops back down slowly to a low of 0.979 from which it slowly starts to rise again.

The first high pressure corresponds with the point at which the balloon first became round and turgid. The lowest pressure occurred when the balloon was about 4 feet in diameter.

Conclusion

The pressure is never more than 0.005 atm greater than ambient pressure. At our normal working size (5 ft diameter), the pressure difference is at its least of 0.002 atm or 0.2% of ambient.

Answer 1: We can assume that the pressure inside the balloon is EQUAL to the outside with only a 0.2% error.

Answer 2: Since the pressure at 5 ft is still well below the maximum, we can feel confident that we are operating a long way from the pop point.

Pressure and the Gas Laws

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Gas Pressure

Gas molecules inside a volume (e.g. a balloon) are constantly moving around freely. During this molecular motion they frequently collide with each other and with the surface of any enclosure there may be (in a small balloon that would be many thousands of billions of collisions each second).
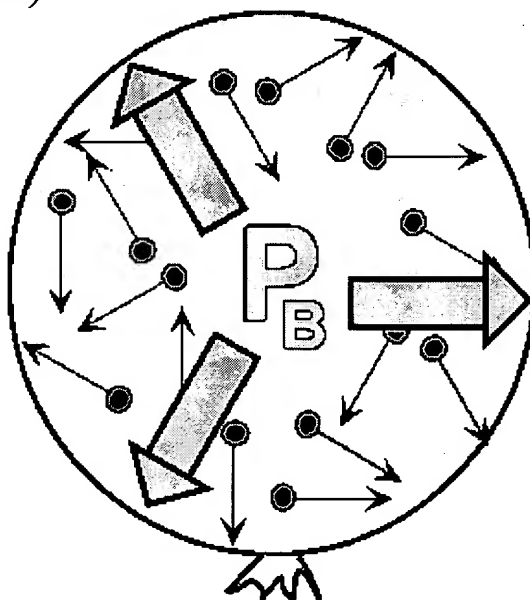
Figure 1: The internal gas pressure in a balloon, P_B, is given by the impacts of moving gas molecules, as they collide with the skin of the balloon from the inside.

The force of impact of a single one such collision is too small to be sensed. However, taken all together, this large number of impacts of gas molecules exerts a considerable force onto the surface of the enclosure: the **gas pressure** (see Figure 1).

The larger the number of collisions per area of enclosure, the larger the pressure:

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}} \quad \text{or} \quad P = \frac{F}{A} \quad \left[\frac{N}{m^2} = Pa\right]_{SI}$$

The SI-unit of pressure is Pascal [Pa], but in Meteorology it is accepted to use millibars [mb], where 100 kPa = 1000 mb.

The *direction* of this gas pressure force is *always perpendicular to the surface* of the enclosure at every point.

Atmospheric Pressure – Pressure Profile

In the example of the balloon (above), there is not only gas inside the balloon (exerting pressure from the inside), but there is also gas (air) on the outside, exerting pressure onto the outside surface of the balloon (Figure 2).
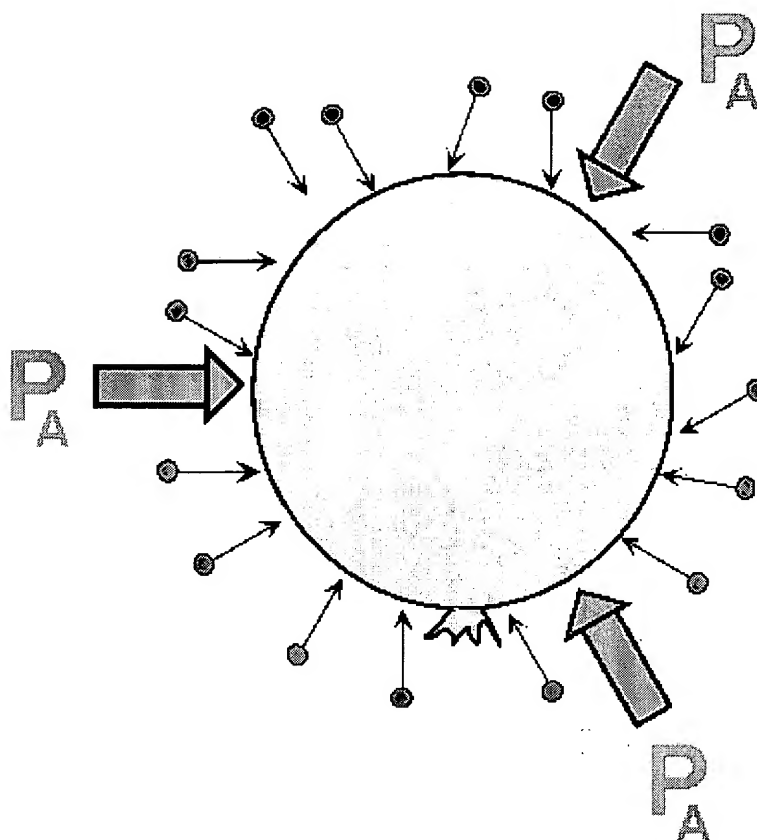
Figure 2: The atmospheric pressure outside a balloon, P_A, is given by the impacts of moving gas molecules, as they collide with the skin of the balloon from the outside.

The rate, at which the skin of the balloon is bombarded by air molecules, is dependent on how tightly the gas molecules are packed, or on the **gas density**:

$$(Gas) Density = \frac{Mass (of\ gas)}{Volume} \quad or \quad \rho = \frac{m}{V} \quad \left[\frac{kg}{m^3}\right]_{SI}$$

Since gas is compressible, its density depends on the force that is used to compress it (see Figure 6-2 in Lutgens and Tarbuck, 7th Ed., 1998). In the atmosphere, the force that compresses the air at the surface is just the weight of all the air in the atmospheric column above it (Figure 3).

$$Atmospheric\ Pressure = \frac{Weight\ of\ Overlying\ Air}{Unit\ Area}$$

At the surface the atmospheric pressure is on average $P_0 = 101.3$ kPa $= 1013$ mb
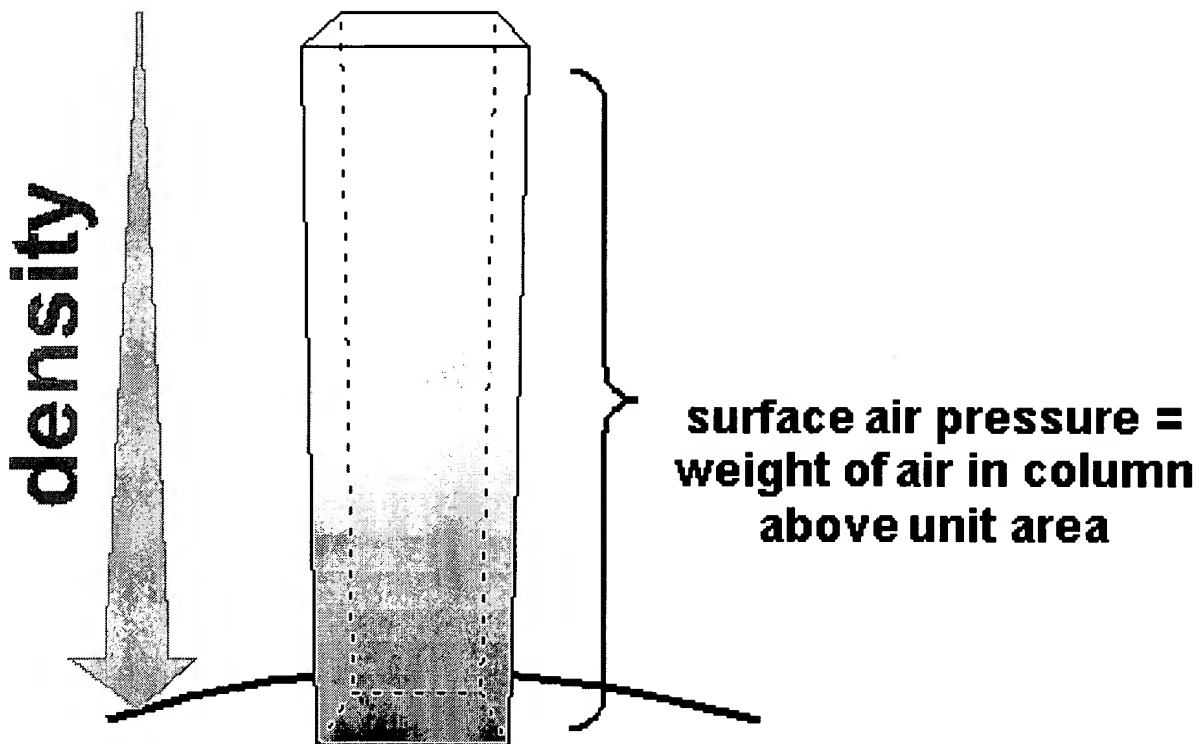
Figure 3: Surface pressure and density in an air column

The higher we go in the atmosphere, the less air remains in the column above us. Thus, the atmospheric pressure always decreases with height. Similarly, air density decreases with height, because the overload to compress the air gets less and less, as we go higher.

Clearly, at the surface, the air is densest. Thus, as we rise from the surface through the first kilometer of the atmosphere, we leave a lot of dense air below us: the overload (and thus the density and pressure) decreases quickly. As a result, the air in the next km is already much less dense than below, and the overload (density and pressure) decrease more slowly. This ever slower decrease of the overload, of density and pressure, is continued for every consecutive vertical stretch of atmosphere.

In fact, at 6 km height, we already have approximately half of the mass of air below us. Thus, the atmospheric pressure at 6 km height is only half that at the surface (i.e., $P_{6km} \approx 500$ mb). Between 6 and 12 km the overload (and the pressure) is halved again: only ¼ of the

surface pressure is left. This consecutive halving gives rise to the following (approximate) pressure profile, starting at a surface pressure of $P_0 = 1000$ mb:

height (km)	Pressure (mb)	Fraction of P_0
0	1000 (= P_0)	1
6	500	1/2
12	250	1/4
18	125	1/8
24	63	1/16
...	...	...

The shape of this profile-curve is called a **negative exponential decrease** (compare Figures 1.12, Lutgens and Tarbuck, 7th Ed., 1998). The density profile behaves in a similar way.

$$\frac{P(z)}{P_0} = \exp\{-\alpha(z - z_0)\}$$

where α is a temperature dependent parameter.

Hydrostatic Pressure – Mercury Barometer

The principle that the pressure at a given level is equivalent to the weight of the overlying column is not only true for air, but for **fluids** (gases and liquids) in general. The pressure generated by an overlying column of fluid is thus termed the **hydrostatic pressure.**

The upper boundary of the air column that gives rise to atmospheric pressure is the vacuum of space. Being rather light, the mass of a column of air with a 1 cm^2 cross section is almost exactly 1 kg. If a much heavier liquid substance is used to balance this air column, only

a relatively small length would be needed. In addition, because the density of liquids does not change with height (most liquids are incompressible), such an equivalent liquid column has a well defined upper boundary (below a vacuum),

One of the heaviest liquids at room temperature is **mercury** (Hg) and the height of the Hg-column that is equivalent to normal pressure (1013 mb) is only **760 mm** long (29.92 "). For this reason, columns of mercury, "hanging" in an inverted vacuum tube, can be used as practical instruments to measure atmospheric pressure (see Figure 6-4, Lutgens and Tarbuck, 1998).

If water were used instead of mercury, the height of the column equivalent to normal pressure would be 10.33 m - not a very practical length of tube to work with.

The Gas Laws

The example of the gas-filled balloon can also be used to explore the basic gas laws (see also Appendix D, p. 414, Lutgens and Tarbuck, 1998).

In the following, lets assume that the balloon is tight, so that the amount or mass of air in it stays the same: m_a = **const**. With density being the ratio of mass per volume, the gas density of the balloon thus varies only with its volume (when mass is held constant).

If we squeeze the balloon, we compress the air and two things will happen:

- the air pressure in the balloon will increase.
- the density of the air in the balloon will increase.

Since density is mass over volume, and the mass stays constant, the

rise in density means that the volume of the balloon decreases: pressure goes up; volume goes down. This finding is expressed more precisely by

Boyle's Law (after Robert Boyle, Irish scientist, around 1600)

Boyle's law states that, *at a constant temperature, the volume of a given mass of gas varies inversely with pressure.* For two states of pressure (P_1, P_2) and two corresponding volumes (V_1, V_2), this is stated mathematically:

$$P_1 \cdot V_1 = P_2 \cdot V_2$$

If $P_2 = 2 \cdot P_1$, it follows that $V_2 = 1/2\ V_1$.

In Lab 3 (see Lab Manual) we are performing another experiment with a balloon that illustrates the principle expressed in

Charles's Law (after Jacques Charles, French scientist, around 1790)

By warming the balloon up, we increase the speed of the moving gas molecules inside it. This in turn increases the rate at which the gas molecules bombard the skin of the balloon. Because the balloon's skin is elastic, it expands upon this increased pushing from inside, and the volume taken up by the same mass of gas increases with temperature. In consequence, the density [=mass/volume] decreases with rising temperature. Cooling the balloon down again will make the balloon shrink.

Thus Charles's law states that *at a constant pressure, the volume of a given mass of gas is directly proportional to its (absolute) temperature.* For two states with temperatures (T_1, T_2) and two corresponding volumes (V_1, V_2):

$$\frac{V_1}{V_2} = \frac{T_1}{T_2} \quad \text{or} \quad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

It must be noted that in this case (and *whenever temperature appears in a multiplication or a division*) the **absolute or Kelvin scale must be used** for temperature.

The Ideal Gas Law or Equation of State

The example used to illustrate Charles's law probably does not follow Charles's law exactly. It is very likely that, during the heating process, when the rate of collisions by the gas molecules increased, the pressure increased as well as the volume. Thus, in practical situations all three variables involved in Boyle's and Charles's law are linked and both principles are in action at the same time:

Pressure: P

Temperature: T

Density: $\rho = m/V$

These variables describe the state of the gas at any one time and are combined in the single relationship known as the ideal gas law or the equation of state:

$$P \cdot V = m \cdot R \cdot T \text{ or } P = \rho \cdot R \cdot T$$

where R, the constant of proportionality, is known as the universal gas constant ($= 287 \text{ J·kg}^{-1}\text{·K}^{-1}$)

When mass and pressure are held constant, the gas law collapses to Charles's law. When mass and temperature are held constant, it is equivalent to Boyle's law. Thus, the gas law combines the two laws.

The equation of state or ideal gas law is one of the most fundamental relationships linking the three variables temperature, pressure and density that describe the thermodynamic state of the atmosphere. Some application of the gas law is involved in most practical problems in meteorology.

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